Periodic Trends

Periodic trends are certain patterns that describe specific aspects of the elements in the periodic table, such as size and properties with electrons. The main periodic trends include: electronegativity, ionization energy, electron affinity, atomic radius, melting point, and metallic character. The periodic trends that arise from the arrangement of the periodic table provide chemists with an invaluable tool to quickly predict an element's properties. These trends exist because of the similar atomic structure of the elements within their respective group families or period and the periodic nature of the elements.

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Periodic Trends for Electronegativity

Electronegativity is a chemical property that describes an atom's ability of attracting and binding to electrons. Since electronegativity describes a largely qualitative property, there is no standardized method of calculating electronegativity. However, the scale that most chemists use in quantifying electronegativity is the Pauling Scale, named after the chemist Linus Pauling. The numbers assigned by the Pauling scale are dimensionless due to electronegativity being largely qualitative. Because electronegativity cannot be calculated, for the Chemistry 2 series, students will be given a periodic table with electronegativity values for each element to work and study from. An example is provided below.

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H 2.1																	He —
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne —
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar —
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr —
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe —
Cs 0.7	Ba 0.9	La–Lu 1.1–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn —

Figure 1. Periodic Table of Electronegativity values

Electronegativity measures an atom's strength to attract and form bonds with electrons. This property exists due to the electronic configuration of atoms. Most atoms prefer to fulfilling the octet rule (having the valence, or outer, shell comprise of 8 electrons). Since elements on the left side of the periodic table have less than a half-full valence shell, the energy required to gain electrons is significantly higher compared to the energy required to lose electrons. As a result, the elements on the left side of the periodic table generally lose electrons in forming bonds. Conversely, elements on the right side of the periodic table are more energy-efficient in gaining electrons to create a complete valence shell of 8 electrons. This effectively describes the nature of electronegativity: the more inclined an atom is to gain electrons, the more likely that atom will pull electrons toward itself.

- As you move to the right across a period of elements, electronegativity ***increases***. When the valence shell of an atom is less than half full, it requires less energy to lose an electron than gain one and thus, it is easier to lose an electron. Conversely, when the valence shell is more than half full, it is easier to pull an electron into the valence shell than to donate one.
- As you move down a group, electronegativity ***decreases***. This is because the atomic number increases down a group and thus there is an increased distance between the valence electrons and nucleus, or a greater atomic radius.
- Important exceptions of the above rules include the noble gases, lanthanides, and actinides. The noble gases possess a complete valence shell and do not usually attract electrons. The lanthanides and actinides possess a more complicated chemistry that does not generally follow any trends. Therefore, noble gases, lanthanides, and actinides do not have electronegativity values.
- As for the transition metals, while they have values, there is little variance among them as you move across the period and up and down a group. This is because of their metaliic properties that affect their ability to attract electrons as easily as the other elements.

With these two general trends in mind, we can deduce that the *most electronegative element is* ***fluorine***, which weighs in at a hefty 3.98 Pauling units.

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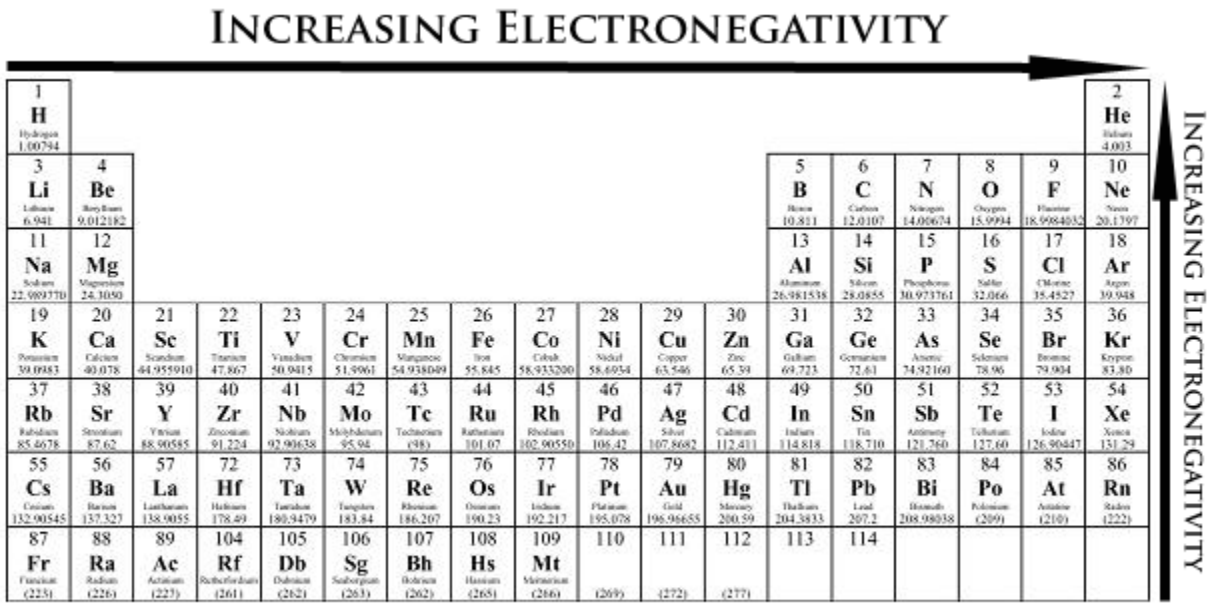

Figure 2. Periodic Table showing Electronegativity Trend

Periodic Trends for Ionization Energy

Ionization Energy is the amount of energy required to remove an electron from a neutral atom in its gaseous phase. Conceptually, ionization energy is considered the opposite of electronegativity. The lower this energy is, the more readily the atom becomes a cation. Therefore, the higher this energy is, the more unlikely the atom becomes a cation. Generally, elements on the right side of the periodic table have a higher ionization energy because their valence shell is nearly filled. Elements on the left side of the periodic table have low ionization energies because of their willingness to lose electrons and become cations. Thus, ionization energy increases from left to right on the periodic table.

Another factor that affects ionization energy is *electron shielding*. Electron shielding describes the ability of an atom's inner electrons to shield its positively-charged nucleus from its valence electrons. When moving to the right on a period of elements, the number of electrons increases and the strength of shielding increases. As a result, it is easier for valence shell electrons to ionize and thus the ionization energy decreases when going down a group. In certain texts, electron shielding may also be known as *screening*.

- The ionization energy of the elements within a period generally increases from left to right. This is due to valence shell stability.
- The ionization energy of the elements within a group generally decreases from top to bottom. This is due to electron shielding.
- The noble gases possess very high ionization energies because of their full valence shell as indicated in the graph. Note that Helium has the highest ionization energy of all the elements.

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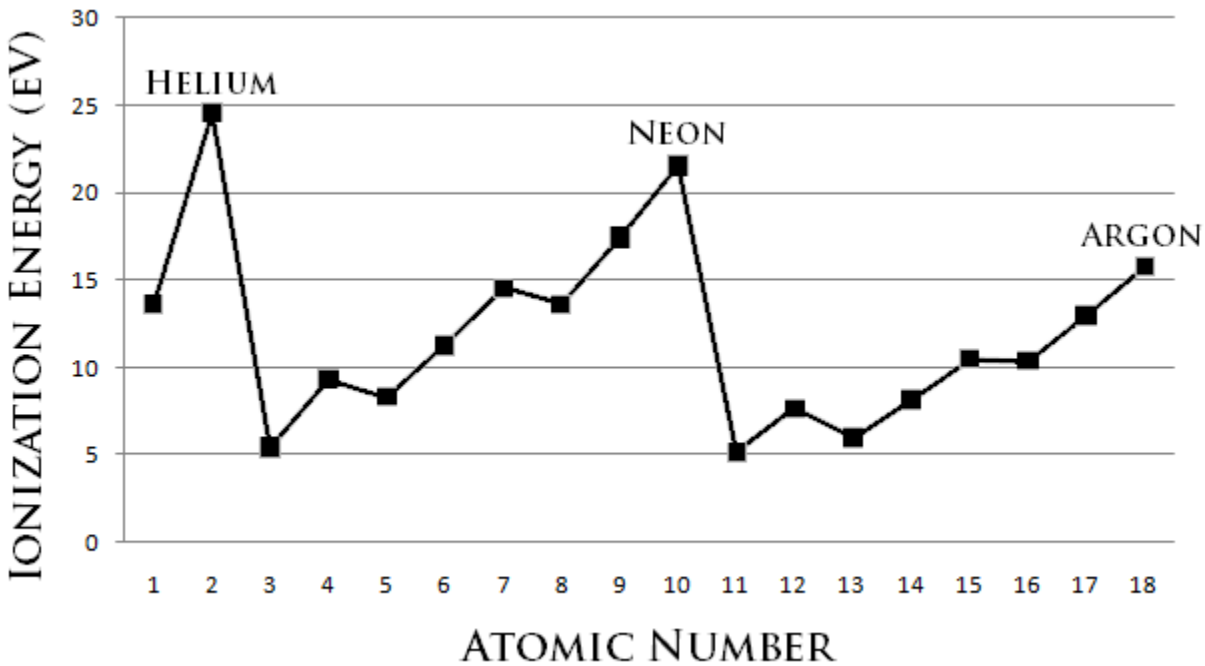

Figure 3. Graph showing the Ionization Energy of the Elements from Hydrogen to Argon

Some elements can have several ionization energies, so we refer to these varying energies as the first ionization energy, the second ionization energy, third ionization energy, etc. The first ionization energy is to the energy needed to remove the outermost, or highest, energy electron and the second ionization energy is the energy required to remove any subsequent high-energy electron from a gaseous cation. Below are the formulas for calculating the first and second ionization energies.

First Ionization Energy:

$$X_{(g)} \rightarrow X^{+}_{(g)} + e^{-}$$

Second Ionization Energy:

$$X^{+}_{(g)} \rightarrow X^{2+}_{(g)} + e^{-}$$

Generally, any subsequent ionization energies (2nd, 3rd, etc.) follow the same periodic trend as the first ionization energy.

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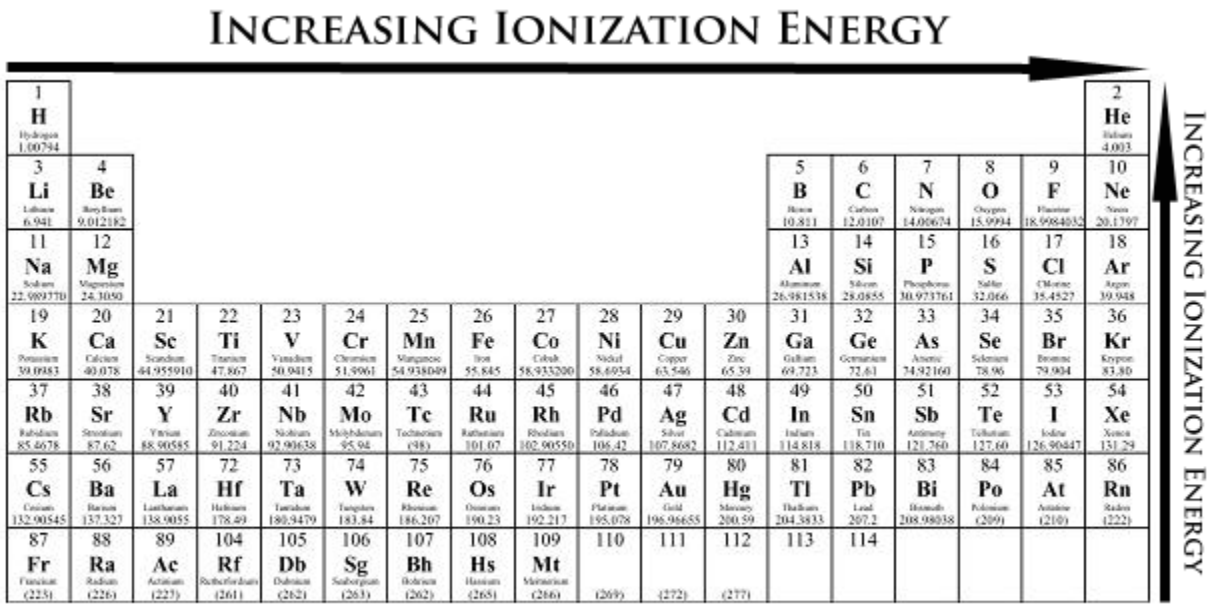

Figure 4. Periodic Table Showing Ionization Energy Trend

Ionization energies decrease as atomic radii increase. This observation is affected by n (the principle quantum number) and Z_{eff} (based on the atomic number and shows how many protons are seen in the atom) on the ionization energy (I). Given by the following equation:

$$I = R_H * Z_{eff}^2/n^2$$

Going across a period, the Z_{eff} *increases* and n (principal quantum number) *remains the same*, so that the ionization energy *increases*.

Going down a group, the n *increases* and Z_{eff} *increases slightly*, the ionization energy *decreases*.

Periodic Trends for Electron Affinity

Like the name suggests, electron affinity describes the ability of an atom to accept an electron. Unlike electronegativity, electron affinity is a quantitative measure that measures the energy change that occurs when an electron is added to a neutral gas atom. When measuring electron affinity, the more negative the value, the more of an affinity to electrons that atom has.

Electron affinity generally decreases down a group of elements because each atom is larger than the atom above it (this is the atomic radius trend, which will be discussed later in this text). This means that an added electron is further away from the atom's nucleus compared to its position in the smaller atom. With a larger distance between the negatively-charged electron and the positively-charged nucleus, the force of attraction is relatively weaker. Therefore, electron affinity decreases. Moving from left to right across a period, atoms become smaller as the forces of attraction become stronger. This causes the electron to move closer to the nucleus, thus increasing the electron affinity from left to right across a period.

- Electron affinity *increases from left to right* within a period. This is caused by the decrease in atomic radius.
- Electron affinity *decreases from top to bottom* within a group. This is caused by the increase in atomic radius.

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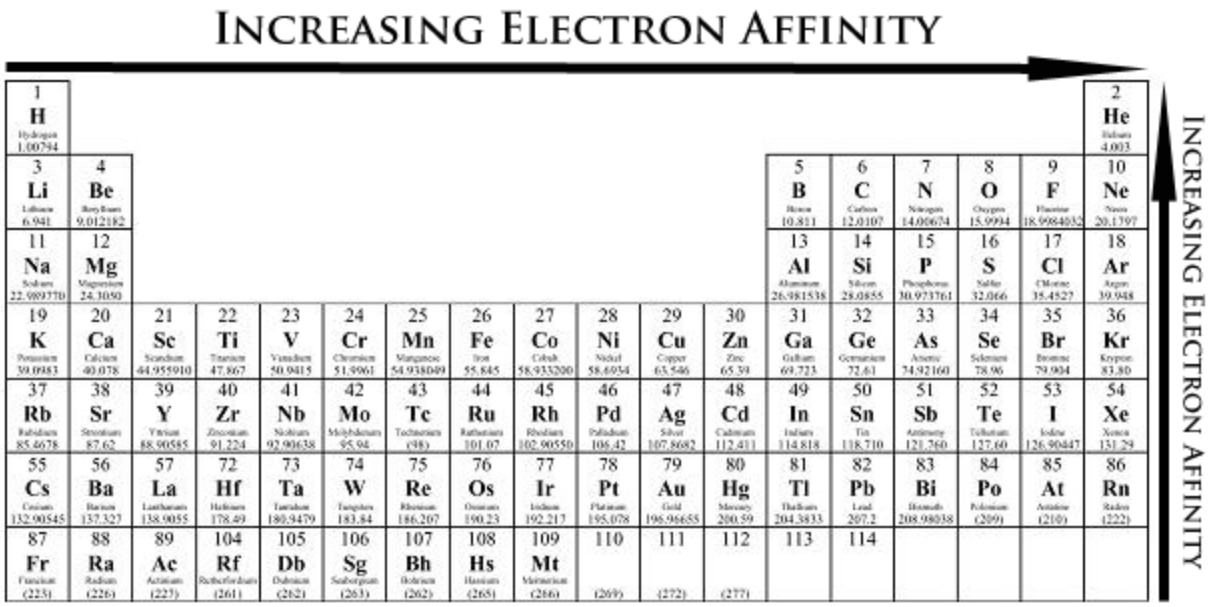

Figure 5. Periodic Table showing Electron Affinity Trend

Periodic Trends for Atomic Radius

For atoms, the atomic radius is one-half the distance between the nuclei of two atoms is (just like a radius is half the diameter of a circle). However, this idea is complicated by the fact that not all atoms are normally bound together in the same way. Some are bound by covalent bonds in molecules, some are attracted to each other in ionic crystals, and others are held in metallic crystals. Nevertheless, it is possible for a vast majority of elements to form covalent molecules in which two like atoms are held together by a single covalent bond. The covalent radius of these molecules is often referred to as the atomic radius. This distance is measured in picometers. Going through each of the elements of the periodic table, patterns of the atomic radius can be seen.

Atomic size gradually decreases from left to right across a period of elements. This is because, within a period or family of elements, all electrons are being added to the same shell. But, at the same time, protons are being added to the nucleus, making it more positively charged. The effect of increasing proton number is greater than that of the increasing electron number; therefore, there is a greater nuclear attraction. This means that the nucleus attracts the electrons more strongly, having the atom's shell pulled closer to the nucleus. The valence electrons are held closer towards the nucleus of the atom. As a result, the atomic radius decreases.

Going *down* a group, it can be seen that atomic radius increases. The valence electrons occupy higher levels due to the increasing quantum number (n). As a result, the valence electrons are further away from the nucleus as the 'n' increases. Electron shielding prevents these outer electrons from being attracted to the nucleus; thus, they are loosely held and the resulting atomic radius is large.

- Atomic radius *decreases from left to right* within a period. This is caused by the increase in the number of protons and electrons across a period. One proton has a greater effect than one electron; thus, a lot of electrons will get pulled towards the nucleus, resulting in a smaller radius.
- Atomic radius *increases from top to bottom* within a group. This is caused by electron shielding.

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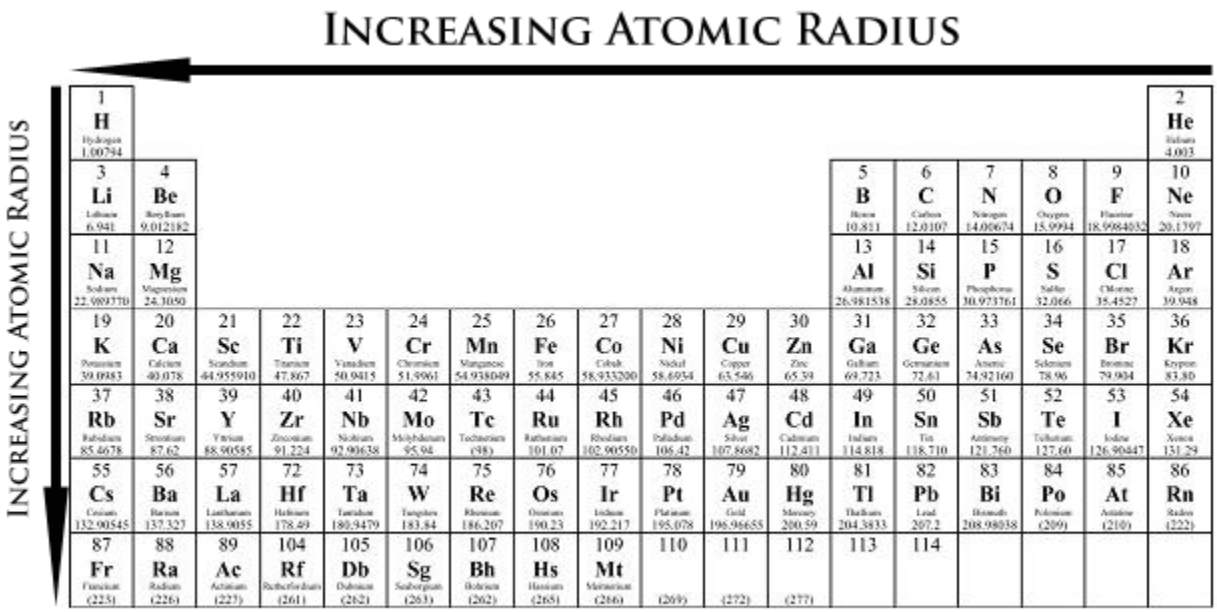

Figure 6. Periodic Table showing Atomic Radius Trend

Periodic Trends for Melting Point

Melting points are the amount of energy required to break a bond(s) to change the solid phase of a substance to a liquid. Generally, the stronger the bonds between the atoms of an element, the higher the energy requirement in breaking that bond. Since temperature is directly proportional to energy, a high bond dissociation energy correlates to a high temperature. Melting points are varied and don't generally form a distinguishable trend across the periodic table. However, certain conclusions can be drawn from the following graph.

- Metals generally possess a *high melting point*.
- Most non-metals possess *low melting points*.
- The non-metal **carbon** possesses *the highest boiling point of all the elements*. The semi-metal boron also possesses a high melting point.

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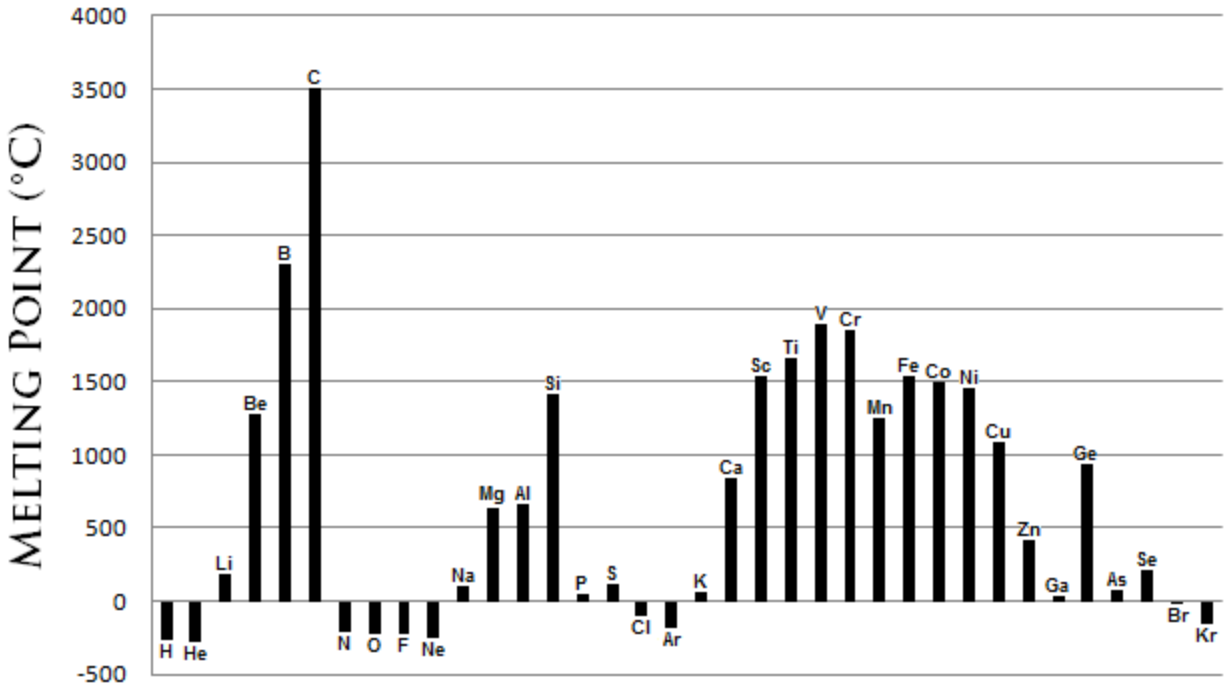

Figure 7. Chart of Melting Points of Various Elements

Periodic Trends for Metallic Character

How easily an atom can lose an electron is a measure of an element's metallic character. As you move from right to left across a period, metallic character increases because the attraction between valence electron and the nucleus is weaker, thus enabling an easier loss of electrons. Metallic character increases as you move down a group because the atomic size is increasing . When the atomic size increases, the outer shells are farther away. The principle quantum number increases and average electron density moves farther from nucleus. The electrons of the valence shell have less of an attraction to the nucleus and, as a result, can lose electrons more readily, causing an increase in metallic character.

- Metallic characteristics decrease from left to right across a period. This is caused by the decrease in radius *(above it is stated that Zeff causes this)*of the atom which allows the outer electrons to ionize more readily.
- Metallic characteristics increase down a group. Electron shielding causes the atomic radius to increase thus the outer electrons ionizes more readily than electrons in smaller atoms.
- Metallic character relates to the ability to lose electrons, and nonmetallic character relates to the ability to gain electrons.
- Another easier way to remember the trend of metallic character is that as you move from left and down towards the bottom-left corner of the periodic table, metallic character increases because you are heading towards Groups 1 and 2, or the Alkali and Alkaline **metal groups**. Likewise, if you move up and to the right to the upper-right corner of the periodic table, metallic character decreases because you are passing by to the right side of the staircase, which indicate the **nonmetals**. These include the Group 8, the **noble gases**, and other common gases such as oxygen and nitrogen.
    - ? In other words:
    - ? Move left across period and down the group: increase metallic character (heading towards alkali and alkaline metals)
    - ? Move right across period and up the group: decrease metallic character (heading towards nonmetals like noble gases)

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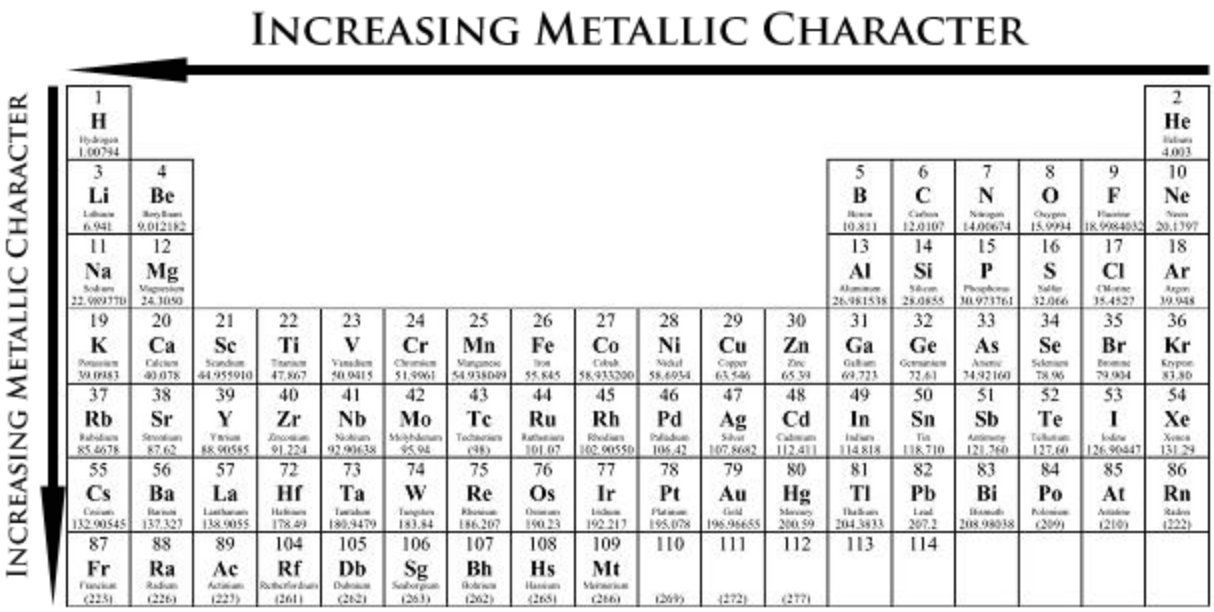

Figure 8. Periodic Table of Metallic Character Trend

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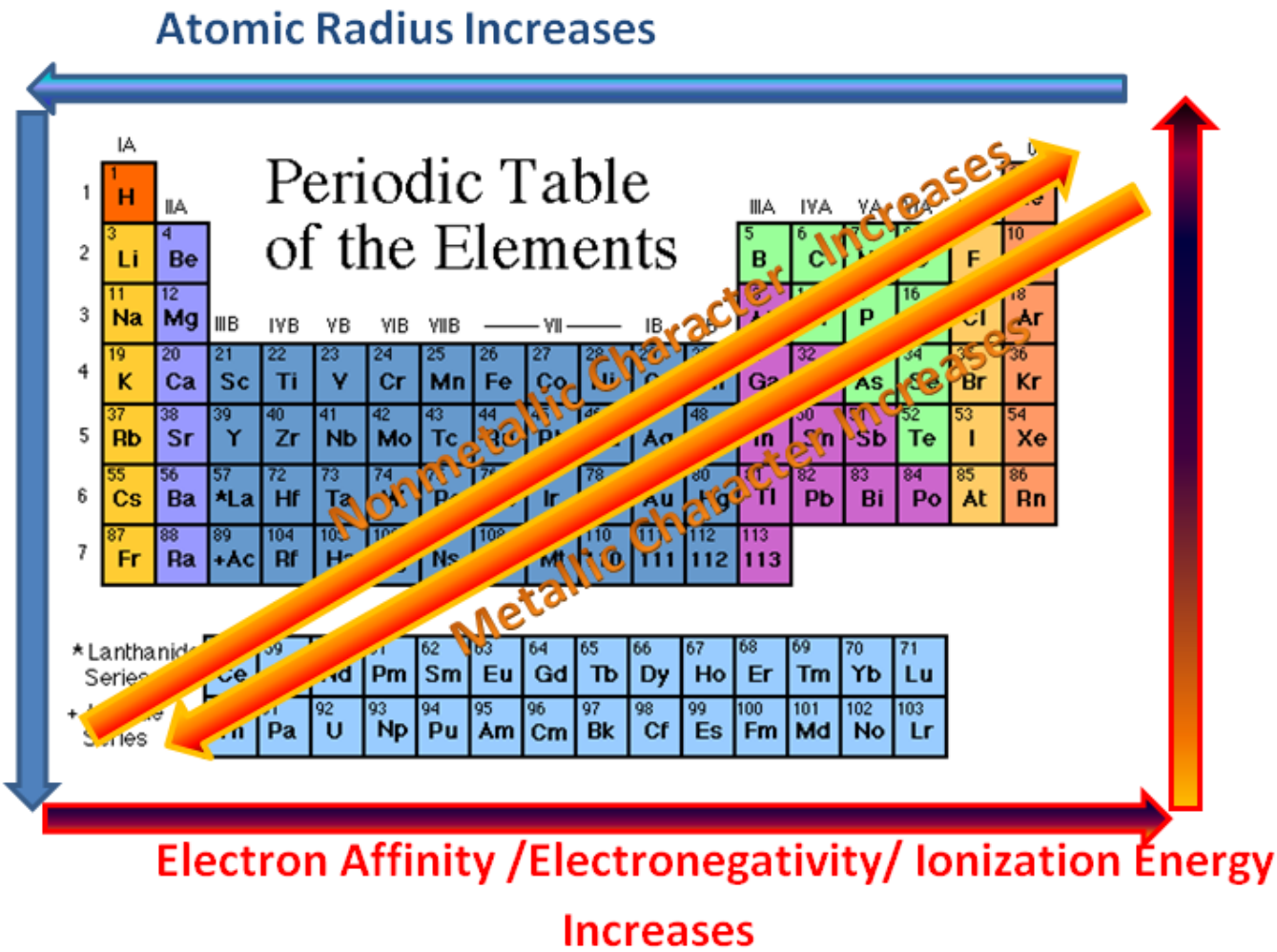
Summary



Outside Links

- http://en.wikipedia.org/wiki/Periodic_trends
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- Ionization Energy Trend: http://www.youtube.com/watch?v=ywqg9PorTAw&feature=youtube_gdata
- Other Periodic Table Trends: http://www.youtube.com/watch?v=XMLd-O6PgVs&feature=youtube_gdata
- Dr. Enderle's (UCD) Lecture on Periodic Trends (2 parts):
    - (Part 1) http://www.youtube.com/watch?v=ZL1uXS43oQY&feature=relmfu
    - (Part 2) http://www.youtube.com/watch?v=p5YeD7jkz30&feature=relmfu

Problems

The following series of problems will review your general understanding of the aforementioned material.

1.) Based on the periodic trends for ionization energy, which do you except to have the highest ionization energy?

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1. A.) Fluorine (F)
2. B.) Nitrogen (N)
3. C.) Helium (He)

2.) Nitrogen has a larger atomic radius than Oxygen.

1. A.) True
2. B.) False

3.) Which do you expect to have more metallic character, Lead (Pb) or Tin (Sn)?

4.) Which element do you expect to have the higher melting point: chlorine (Cl) or bromine (Br)?

5.) Which element do you expect to be more electronegative, sulfur (S) or selenium (Se)?

6) Why is the electronegativity value of most noble gases equal to zero?

7) Arrange the following atoms according to decreasing effective nuclear charge experienced by their valence electrons: S, Mg, Al, Si

8) Rewrite the following list in order of decreasing electron affinity: Fluorine (F), Phosphorous (P), Sulfur (S), Boron (B).

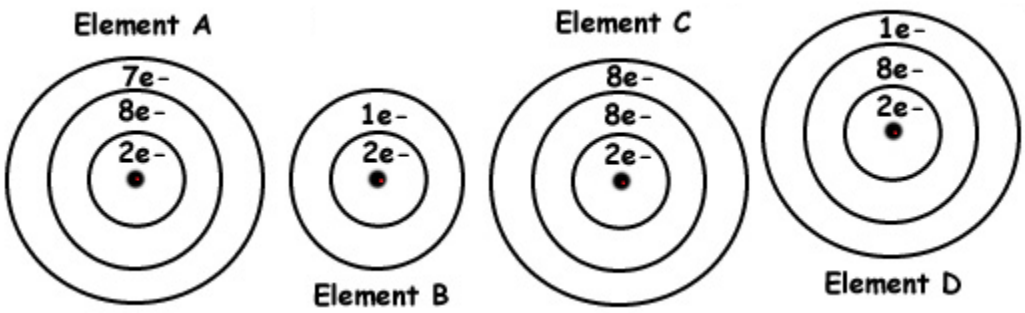
9) An atom with an atomic radius smaller than that of Sulfur (S) is _____.

1. A.) Oxygen (O)
2. B.) Chlorine (Cl)
3. C.) Calcium (Ca)
4. D.) Lithium (Li)
5. E.) None of the above

10) A nonmetal will have a smaller ionic radius when compared to a metal of the same period.

1. A.) True B.) False

11) Which one of the following has the lowest first ionization energy?



1. A. Element A
2. B. Element B
3. C. Element C
4. D. Element D

Solutions

1. Answer: C.) Helium (He)

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Explanation: Helium (He) has the highest ionization energy because, like other noble gases, Helium's valence shell is full. Because of this, Helium is stable and does not readily lose or gain electrons.

2. Answer: A.) True

Explanation: According to periodic trends, atomic radius increases from right to left on the periodic table. Therefore, we would expect Nitrogen to be larger than Oxygen.

3. Answer: Lead (Pb)

Explanation: Lead and Tin share the same column. According to periodic trends, metallic character increases as you go down a column. Lead is underneath Tin therefore we would expect Lead to possess more metallic character.

4. Answer: Bromine (Br)

Explanation: According to periodic trends, in non-metals, melting point increases down a column. Since chlorine and bromine share the same column, we would expect bromine to possess the higher melting point.

5. Answer: Sulfur (S)

Explanation: Note that sulfur and selenium share the same column. Periodic trends tell us that electronegativity increases up a column. This indicates that sulfur is more electronegative than selenium.

6. Answer: Most noble gases have full valence shells.

Explanation: Because of their full valence electron shell, the noble gases are extremely stable and do not readily lose or gain electrons.

7. Answer: S > Si > Al > Mg.

Explanation: The electrons above a closed shell are shielded by the closed shell. S has 6 electrons above a closed shell, so each one feels the pull of 6 prontons in the nucleus.

8. Answer: Fluorine (F)>Sulfur (S)>Phosphorous (P)>Boron (B)

Explanation: According to periodic trends, the electron affinity generally increases from left to right and from bottom to top.

9. Answer: C.) Oxygen (O)

Explanation: Periodic trends indicate that atomic radius increases up a group and from left to right across a period. Therefore, oxygen is expect to have a smaller atomic radius than of sulfur.

10. Answer: B.) False

Explanation: The reasoning behind this lies in understanding that a metal usually loses an electron in becoming an ion while a non-metal gains an electron. This results in a smaller ionic radius for the metal ion and a larger ionic radius for the non-metal ion.

11. Element D

Explanation: Element A, B and D have the same number of elentrons in the inner shell, but element D has the least number of eletrons in the outer shell which requires the lowest ionization energy.

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